

Name Key

Equilibrium Practice Test

1. Consider the reaction system, $\text{CoO(s)} + \text{H}_2\text{(g)} \rightleftharpoons \text{Co(s)} + \text{H}_2\text{O(g)}$.

The equilibrium constant expression is

- a) $\frac{[\text{CoO}][\text{H}_2]}{[\text{Co}][\text{H}_2\text{O}]}$ b) $\frac{[\text{Co}][\text{H}_2\text{O}]}{[\text{CoO}][\text{H}_2]}$ c) $\frac{[\text{Co}][\text{H}_2\text{O}]}{[\text{H}_2]}$
 d) $\frac{[\text{H}_2]}{[\text{H}_2\text{O}]}$ e) $\frac{[\text{H}_2\text{O}]}{[\text{H}_2]}$

2. Given the equilibrium,

$2\text{SO}_2\text{(g)} + \text{O}_2\text{(g)} \rightleftharpoons 2\text{SO}_3\text{(g)}$, if this equilibrium is established by beginning with equal number of moles of SO_2 and O_2 in a 1.0 Liter bulb, then the following **must** be true at equilibrium:

- a) $[\text{SO}_2] = [\text{SO}_3]$ b) $2[\text{SO}_2] = 2[\text{SO}_3]$ c) $[\text{SO}_2] = [\text{O}_2]$
 d) $[\text{SO}_2] < [\text{O}_2]$ e) $[\text{SO}_2] > [\text{O}_2]$

SO₂ reacts at double the rate, so less is remaining at equilibrium.

Questions 3 & 4 refer to the following:

At a given temperature, 0.300 mole NO, 0.200 mol Cl_2 and 0.500 mol NOCl were placed in a 25.0 Liter container. The following equilibrium is established: $2\text{NOCl(g)} \rightleftharpoons 2\text{NO(g)} + \text{Cl}_2\text{(g)}$

3. At equilibrium, 0.600 mol of NOCl was present. The number of **moles** of Cl_2 present at equilibrium is

- a) 0.050 b) 0.100 c) 0.150
 d) 0.200 e) 0.250

	NOCl	NO	Cl ₂
I	.5	.3	.2
C	+.1	+.1	+.05
E	.6	.2	.15

} per 25L

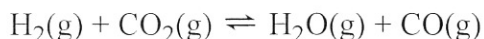
4. The equilibrium constant, K_c , is:

- a) 4.45×10^{-4} b) 6.67×10^{-4} c) 0.111
 d) 0.167 e) 1500

$$K_c = \frac{[\text{NO}]^2 [\text{Cl}_2]}{[\text{NOCl}]^2}$$

$$K_c = 0.0006$$

5. At 985°C, the equilibrium constant for the reaction,



is 1.63. What is the equilibrium constant for the reverse reaction?

- a) 1.63 b) 0.815 c) 2.66
 d) 0.613 e) 1.00

$$K_{\text{rev}} = \frac{1}{K_{\text{forward}}} = \frac{1}{1.63} = .613$$

equal moles of gas

6. What is the relationship between K_p and K_c for the reaction, $2\text{ICl(g)} \rightleftharpoons \text{I}_2\text{(g)} + \text{Cl}_2\text{(g)}$?

- a) $K_p = K_c(RT)^{-1}$ b) $K_p = K_c(RT)$ c) $K_p = K_c(RT)^2$
 d) $K_p = K_c$ e) $K_p = K_c(2RT)$

7. For the reaction $2\text{NO}_2\text{(g)} \rightleftharpoons \text{N}_2\text{O}_4\text{(g)}$, K_p at 25°C is 7.3, when all partial pressures are expressed in atmospheres. What is K_c for this reaction? [$R=0.0821 \text{ L}\cdot\text{atm}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$]

- a) 4270 b) 0.0119 c) 0.291
 d) 179 e) 2.06

$K_p = K_c(RT)^{-1}$
 $7.3 = K_c(0.0821 \times 298)^{-1}$
 $K_c = 179$

8. 0.200 mol NO is placed in a one liter flask at 2273 K. After equilibrium is attained, 0.0863 mol N_2 and 0.0863 mol O_2 are present. What is K_c for this reaction?



- a) 9.92 b) 3.15 c) 0.0372
 d) 39.7 e) 0.576

	NO	N ₂	O ₂
I	.2	0	0
C	-.1726	+.0863	+.0863
E	.0274	.0863	.0863

$$K_c = \frac{[.0863][.0863]}{[.0274]^2} = 9.92$$

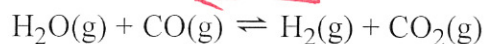
9. $\text{N}_2\text{O}_4\text{(g)} \rightleftharpoons 2\text{NO}_2\text{(g)}$

At 25°C , 0.11 mole of N_2O_4 reacts to form 0.10 mole of N_2O_4 and 0.02 mole of NO_2 . At 90°C , 0.11 mole of N_2O_4 forms 0.050 mole of N_2O_4 and 0.12 mole of NO_2 . From these data we can conclude

- a) N_2O_4 molecules react by a second order rate law.
 b) N_2O_4 molecules react by a first order rate law.
 c) the reaction is exothermic.
 d) N_2O_4 molecules react faster at 25°C than at 90°C .
 e) the equilibrium constant for the reaction above increases with an increase in temperature.

Impossible to tell

10. For the equilibrium system



$$\Delta H = -42 \text{ kJ/mol (exothermic)}$$

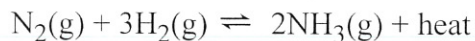
K_c equals 0.62 at 1260 K. If 0.10 mole each of H_2O , CO , H_2 and CO_2 (each at 1260 K) were placed in a 1.0-Liter flask at 1260 K, when the system came to equilibrium...

The temperature would The mass of CO would

- a) decrease increase
 b) decrease decrease
 c) remain constant increase
 d) increase decrease
 e) increase increase

$Q = \frac{[.10][.10]}{[.10][.10]} = 1$
 $Q > K_c$ (will shift left)

11. For the reaction system,



the conditions that would favor maximum conversion of the reactants to products would be

- a) high temperature and high pressure
 - b) high temperature, pressure unimportant
 - c) high temperature and low pressure
 - ☒ d) low temperature and high pressure
 - e) low temperature and low pressure
12. Solid HgO, liquid Hg, and gaseous O₂ are placed in a glass bulb and are allowed to reach equilibrium at a given temperature.



endothermic
 $\Delta H = +43.4 \text{ kcal}$

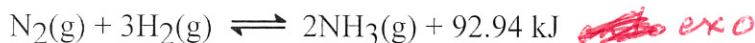
The mass of HgO in the bulb could be increased by

- a) adding more Hg.
- b) removing some O₂.
- ☒ c) reducing the volume of the bulb.
- d) increasing the temperature.
- e) removing some Hg.

$K_c = [\text{O}_2]$ (heterogeneous)

Hg doesn't have a concentration in the gas phase.

13. Consider the following system at equilibrium:



~~endo~~ *exo*

Which of the following changes will shift the equilibrium to the right?

- ~~I.~~ increasing the temperature
- ☒ II. decreasing the temperature
- ~~III.~~ increasing the volume
- ☒ IV. decreasing the volume
- ~~V.~~ removing some NH₃
- ~~VI.~~ adding some NH₃
- ~~VII.~~ removing some N₂
- ☒ VIII. adding some N₂

- a) I, IV, VI, VII
- b) II, III, V, VIII
- c) I, VI, VIII
- d) I, III, V, VII
- ☒ e) II, IV, V, VIII

14. Which of the following statements concerning equilibrium is not true?

- a) A system that is disturbed from an equilibrium condition responds in a manner to restore equilibrium.
- b) Equilibrium in molecular systems is dynamic, with two opposing processes balancing one another.
- c) The value of the equilibrium constant for a given reaction mixture is the same regardless of the direction from which equilibrium is attained.
- d) A system moves spontaneously toward a state of equilibrium.
- ☒ e) The equilibrium constant is independent of temperature.

15–17. The following questions refer to the equilibrium shown here:



15. What would happen to the system if oxygen were added?

- ~~a)~~ More ammonia would be produced.
- ~~b)~~ More oxygen would be produced.
- ☒ c) The equilibrium would shift to the right.
- d) The equilibrium would shift to the left.
- e) Nothing would happen.

16. What would happen to the system if the pressure were decreased?

- a) Nothing would happen.
- b) More oxygen would be produced.
- c) The water vapor would become liquid water.
- d) The ammonia concentration would increase.
- ☒ e) The NO concentration would increase.

shifts right (more gas)

17. For a certain reaction at 25.0°C, the value of K is 1.2×10^{-3} . At 50.0°C the value of K is 3.4×10^{-1} . This means that the reaction is

- a) exothermic.
- ☒ b) endothermic.
- c) never favorable.
- d) More information is needed.
- e) None of these (a-d)

*As T increases, K increases
(shifts right)*

Answers: 1E 2D 3C 4B 5D 6D 7D 8A 9E 10A 11D 12C 13E 14E 15C 16E 17B